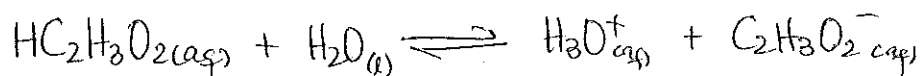


CHM129

Acid-Base Equilibrium: Buffers

1. Calculate the pH of a buffer solution that is 0.100 M $\text{HC}_2\text{H}_3\text{O}_2$ and 0.100 M $\text{NaC}_2\text{H}_3\text{O}_2$. $K_a = 1.8 \times 10^{-5}$



	$[\text{HC}_2\text{H}_3\text{O}_2]$	$[\text{H}_3\text{O}^+]$	$[\text{C}_2\text{H}_3\text{O}_2^-]$
initial	0.100	0.00	0.100
change	-x	+x	+x
equil.	0.100-x	x	0.100+x

$$\frac{[\text{HA}]_{\text{ini}}}{K_a} = \frac{0.100}{1.8 \times 10^{-5}} = 5.6 \times 10^3 > 400$$

Assume x is small

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = 1.8 \times 10^{-5}$$

$$\frac{(x)(0.100+x)}{(0.100-x)} = 1.8 \times 10^{-5}$$

$$\frac{(x)(0.100)}{0.100} = 1.8 \times 10^{-5}$$

$$x = 1.8 \times 10^{-5} \text{ M} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(1.8 \times 10^{-5})$$

$$\text{pH} = \underline{\underline{4.74}}$$

2. Find the pH, using the Henderson-Hasselbalch equation, of a buffer solution that is 0.100 M $\text{HC}_2\text{H}_3\text{O}_2$ and 0.100 M $\text{NaC}_2\text{H}_3\text{O}_2$. $K_a = 1.8 \times 10^{-5}$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\text{pH} = -\log(1.8 \times 10^{-5}) + \log \frac{(0.100)}{(0.100)}$$

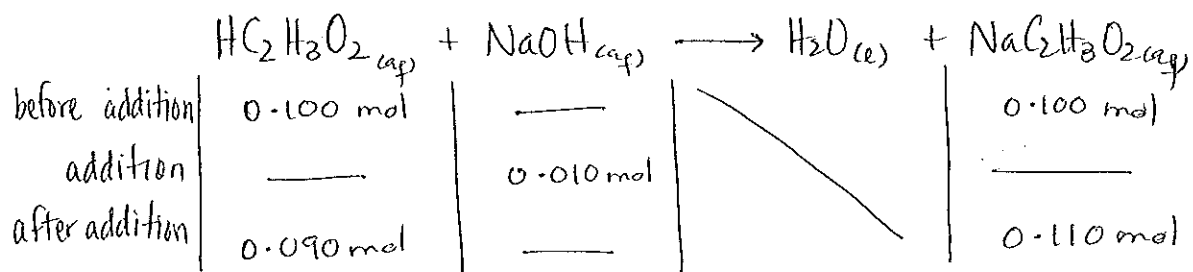
$$\text{pH} = \text{p}K_a = \underline{\underline{4.74}}$$

3. A 1.0 L buffer solution contains 0.100 mol $\text{HC}_2\text{H}_3\text{O}_2$ and 0.100 mol $\text{NaC}_2\text{H}_3\text{O}_2$. It has an initial $\text{pH} = 4.74$. The value of K_a for $\text{HC}_2\text{H}_3\text{O}_2$ is 1.8×10^{-5} .

(a) Calculate the new pH after adding 0.010 mol of solid NaOH to the buffer.

(b) For comparison, compute the pH after adding 0.010 mol of solid NaOH to 1.0 L of pure water.

* Ignore any small changes in volume due to the addition of NaOH .



$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\text{pH} = -\log(1.8 \times 10^{-5}) + \log \frac{(0.110)}{(0.090)} = \underline{\underline{4.83}}$$

$$2) [\text{NaOH}] = \frac{0.010 \text{ mol}}{1.0 \text{ L}} = 0.010 \text{ M} = [\text{OH}^-]$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(0.010) = 2.00$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 2.00 = \underline{\underline{12.00}}$$